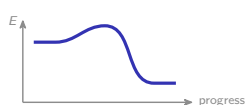


Thermochemistry

AP Chemistry • Unit 6

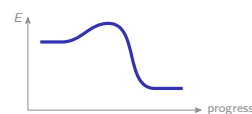
AP Chemistry



Thermochemistry

What we will cover

- 1 Endothermic and exothermic
- 2 Energy diagrams and ΔH
- 3 Calorimetry
- 4 Phase changes
- 5 Bond enthalpies
- 6 Enthalpy of formation
- 7 Hess's law



1

Endo/exo

Thermochemistry

↳ Endo/exo

Endothermic and exothermic

Energy flows between the **system** 系统和 its **surroundings** 环境:

exothermic 放热: releases energy (warms)

endothermic 吸热: absorbs energy (cools)



Endothermic packs absorb heat from surroundings

2

Energy diagrams

Thermochemistry

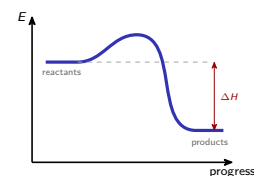
↳ Energy diagrams

Energy diagrams and enthalpy

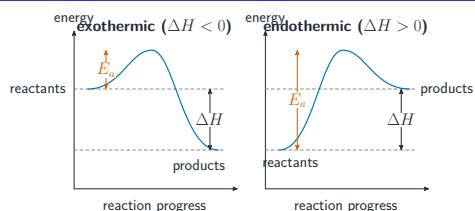
The **enthalpy change** 焓变 ΔH is the heat at constant pressure:

- products **lower** $\Rightarrow \Delta H < 0$ (exothermic)
- products **higher** $\Rightarrow \Delta H > 0$ (endothermic)

ΔH is **extensive** 广延量 – double the moles, double ΔH .



Energy Diagrams



Energy diagrams: exothermic products sit below the reactants, endothermic above

3

Calorimetry

Calorimetry

Specific heat 比热容 c warms 1 g by 1° :

Heat transferred

$$q = mc\Delta T$$

Calorimetry 量热法: an isolated system, so $q_{\text{rxn}} = -q_{\text{water}}$

Worked example

100 g water warms 5.0° :

$$q = 100(4.18)(5.0) = 2090 \text{ J}$$

The reaction released -2090 J – exothermic. For 0.010 mol: $\Delta H = -209 \text{ kJ/mol}$.

4

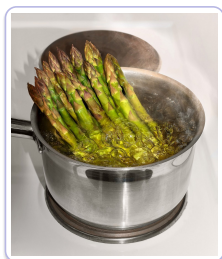
Phase changes

Energy of phase changes

A **phase change** 相变 overcomes **intermolecular forces** 分子间作用力, not bonds.

On a **heating curve** 加热曲线, temperature stays **constant** during a phase change – energy goes into breaking IMFs.

Heat of vaporization 汽化热 > **heat of fusion** 融化热 (boiling separates molecules fully).



Phase changes absorb or release latent heat

5

Bond enthalpies

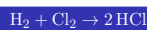
Bond enthalpies

Estimate ΔH from the bonds that change:

Broken minus formed

$$\Delta H \approx \sum(\text{broken}) - \sum(\text{formed})$$

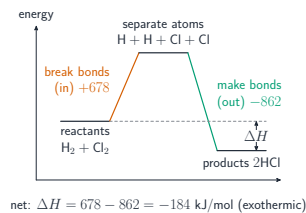
Breaking is endothermic (in), forming exothermic (out).



break H-H (436) + Cl-Cl (243); form 2 H-Cl (431 each):

$$(679) - (862) = -183 \text{ kJ}$$

Breaking bonds takes in energy; making bonds



Breaking bonds takes in energy; making bonds releases it

6 Formation & Hess

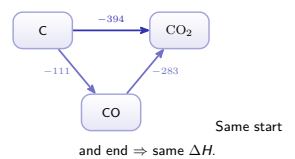
Enthalpy of formation and Hess's law

From a table of ΔH_f°

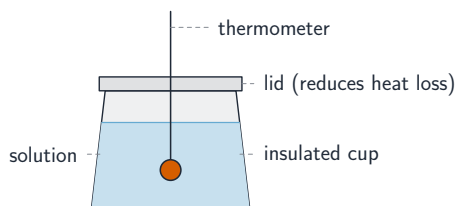
$$\Delta H_{\text{rxn}}^\circ = \sum \Delta H_f^\circ(\text{prod}) - \sum \Delta H_f^\circ(\text{react})$$

An element in its standard state has $\Delta H_f^\circ = 0$.

Hess's law 赫斯定律: ΔH is a **state function** 状态函数 – add steps. **Reverse** flips the sign; **scale** multiplies it.



Heat Capacity and Calorimetry



Calorimetry: measure the temperature change of a known mass of solution

Heat, equilibrium, and the enthalpies

heat 热量	flows from hot to cold until objects reach thermal equilibrium 热平衡
equilibrium means	equal temperature – not equal energy, and not no motion
enthalpy of reaction 反应焓	ΔH_{rxn} – the heat released or absorbed at constant pressure
the sign	negative for exothermic, positive for endothermic
standard enthalpy of formation 标准生成焓	ΔH_f° – one mole of a compound from its elements in their standard states
why it is useful	$\Delta H_{\text{rxn}} = \sum \Delta H_f^\circ(\text{products}) - \sum \Delta H_f^\circ(\text{reactants})$

An element in its standard state has $\Delta H_f^\circ = 0$ by definition. Forgetting that is what breaks the subtraction.

Enthalpy of Formation

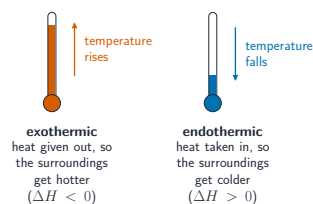
formation ΔH_f :

elements → 1 mol of the compound

combustion ΔH_c :1 mol substance + O₂ → CO₂ + H₂O

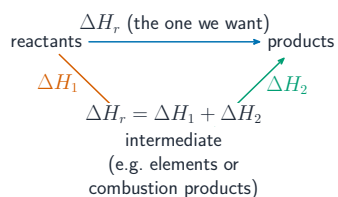
Formation makes a compound from its elements; combustion burns it in oxygen

Endothermic and Exothermic Processes



An exothermic reaction warms the surroundings; an endothermic one cools them

Hess's Law



Hess's law: the direct and indirect routes give the same total enthalpy change

Exam tips

- An **exothermic** reaction has $\Delta H < 0$ (feels hot); endothermic has $\Delta H > 0$ – always give ΔH a sign and units.
- In calorimetry use $q = mc\Delta T$ with the mass of the **water/solution**; the reaction releases what the water gains.
- Bond enthalpies: $\Delta H \approx \sum(\text{bonds broken}) - \sum(\text{bonds made})$ – breaking is endothermic, making is exothermic (the classic sign trap).
- **Hess's law**: the total ΔH is the sum of the steps' – reverse a step and flip the sign, scale a step and scale ΔH .
- During a **phase change** the temperature stays constant while energy goes into the forces between particles.

7 Recap

Unit 6 – key takeaways

1 Enthalpy

 $\Delta H < 0$ exothermic; break absorbs, form releases

2 Calorimetry

 $q = mc\Delta T$; $q_{\text{rxn}} = -q_{\text{water}}$

3 Phase change

constant T ; vaporization > fusion

4 Enthalpy paths

bond enthalpies; ΔH_f° ; Hess adds steps

Check yourself

- 1 If products sit lower than reactants, what is the sign of ΔH ?
- 2 Does breaking a bond absorb or release energy?
- 3 What happens to temperature during a phase change?
- 4 Reverse a reaction – what happens to its ΔH ?



Answers 1. negative 2. absorb 3. it stays constant 4. its sign flips